Unit #5
Polyatomic Ions / Mole / Molarity / Electrochemistry
Instructor's notes

Introduction

1) Remember how to write / draw:
   a) Electron Configuration
   b) Orbital Notation
   c) Electron Dot Notation
   d) Octet Rule

2) **Ion-** an atom or group of atoms that has either lost of gained electron(s) and as a result has an electric charge.
   There are two basic types of ions.
   a) **Cation-** is a positively charged ion.
   b) **Anion-** is a negatively charged ion.

3) We will draw single atom ions exactly like atoms with two exceptions
   a) If the atom has at least 5, but less than 8 valence electrons; we will draw as many additional electrons as needed to acquire a total of 8 valence electrons. (This will give the ion an Octet of valence electrons.)
      i) After the "atomic symbol" of the ion, in the upper right corner of the symbol, we will write a negative sign (or digit, if more than one) for each additional electron we added to achieve the octet of valence electrons. **What type of ion would the following be ?** *Answer: Anions.*
         (1) Example 1:
            Nitrogen has how many valence electrons ? *Answer: 5*
            How many additional electrons must be added to achieve the octet of valence electrons ? *Answer: 3*
            How would we write the symbol for a Nitrogen ion ?
            Answer: N^{-3}
         (2) Example 2:
            Bromine has how many valence electrons ? *Answer: 7*
            How many additional electrons must be added to achieve the octet of valence electrons ? *Answer: 1*
            How would we write the symbol for a Bromine ion ?
            Answer: Br^{-1}

(3) Example 3:

Oxygen has how many valence electrons ? *Answer: 6*
How many additional electrons must be added to achieve the octet of valence electrons ? *Answer: 2*
How would we write the symbol for a Oxygen ion ? *Answer: O^{-2}*

b) If the atom has less than 4 valence electrons; we will remove as many electrons as needed to move back to the nearest noble gas core. A total of 8 valence electrons. (This will give the ion an Octet of valence electrons.)

i) After the "atomic symbol" of the ion, in the upper right corner of the symbol, we will write a positive sign (or digit, if more than one) for each additional electron we subtracted to achieve the nearest noble gas core. An octet of valence electrons. **What type of ion would the following be ?** *Answer: Cation.*

(1) Example 1:

Sodium has how many valence electrons ? *Answer: 1*
How many additional electrons must be subtracted to achieve the octet of valence electrons ? *Answer: 1*
How would we write the symbol for a Sodium ion ? *Answer: Na^{+1}*

(2) Example 2:

Aluminum has how many valence electrons ? *Answer: 3*
How many additional electrons must be subtracted to achieve the octet of valence electrons ? *Answer: 3*
How would we write the symbol for a Sodium ion ? *Answer: Al^{+3}*

(3) Example 3:

Barium has how many valence electrons ? *Answer: 2*
How many additional electrons must be subtracted to achieve the octet of valence electrons ? *Answer: 2*
How would we write the symbol for a Sodium ion ? *Answer: Ba^{+2}*

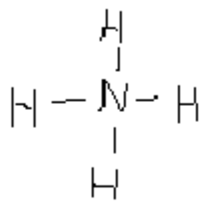
4) **Polyatomic ion-** a type of ion made of more than 2 atoms. These will also be classified as either an ion or cation.
   a) 4 steps to drawing polyatomic ions.
      i) Determine and write down the "4 key numbers". These are:
         (1) **# of protons** (add up the atomic number of all involved atoms).
         (2) **# of total electrons** (add up the atomic numbers (which equals the number of total electrons in a neutral atom) of all atoms in the ion and:
            (a) if the charge is negative, add the value of that charge to the total electrons.
            (b) If the charge is positive, subtract the value of that charge from the total electrons.
         (3) **# of core electrons** (atomic number minus the number of "s" and "p" electrons in the highest energy level).
         (4) **# of valence electrons** (determined by subtracting the # of core electrons from the # of total electrons).
      ii) Draw the "skeleton" of the ion (See the example below)
      iii) Give all atoms eight electrons (except Hydrogen, it only wants two), (remember each bond (represented by a dash) counts as 2 electrons).
      iv) Count the number of electrons drawn and compare this to the calculated number ("Key number #4").
         (1) Adjust as necessary by adding double or triple bonds to reduce the number of electrons while still ensuring all atoms still have an octet of valence electrons.
         (2) Draw brackets and valence charge to finish.

   b) Examples of how to draw polyatomic ions.
      i) NH_4^{+1} (ammonium)

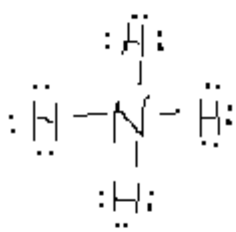
         STEP 1: *Determine "4 key numbers"*

         P = *11*
         E = *10*
         C = *2*
         V = *8*

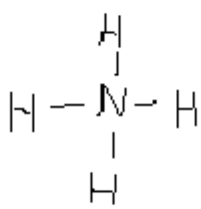
STEP 2: *Draw skeleton*



STEP 3: *Give all atoms eight electrons. EXCEPT: Hydrogen.*

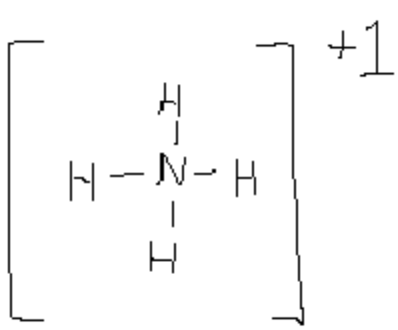


Note this is incorrect. Hydrogen does not want 8 electrons.



This is correct. Nitrogen has 8 electrons. Hydrogen has 2 electrons each.

STEP 4: *Check number of electrons, adjust if necessary, and draw brackets to finish.*

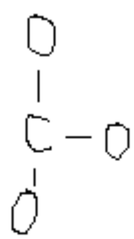


ii) CO_3^{-2} (carbonate)

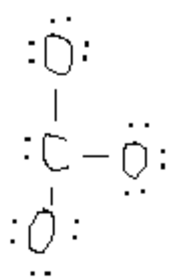
STEP 1: *Determine "4 key numbers"*

P = *30*
E = *32*
C = *8*
V = *24*

STEP 2: *Draw skeleton*

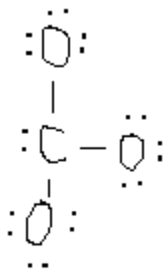


STEP 3: *Give all atoms eight electrons. EXCEPT: Hydrogen.*



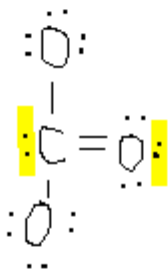
STEP 4: *Check number of electrons, adjust if necessary, and draw brackets to finish.*

A.



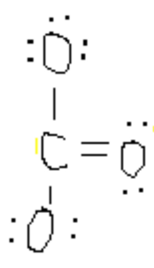
26 electrons are too many. How many should there be ? Answer: *24* We will have to try adding a double bond. Follow along to the next step.

B.



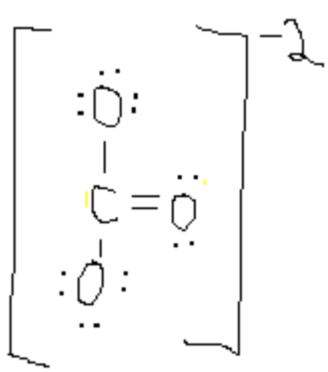
Notice the addition of the double bond, it can be drawn alongside any single bond. However, whenever a double bond is drawn, a pair of electrons must be removed from each of the atoms at the end of that bond. (highlighted yellow electrons). Follow along to the next step.

C.



We now have 24 electrons, which according to our "4 Key Numbers", is the correct number of electrons. Only one more step.

D.

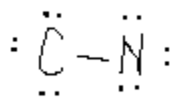


Good job. It is finished correctly.

iii) CN^- (cyanide)

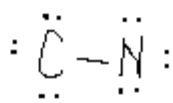
We will assume you have already calculated the "4 Key Numbers". What are there values ? *Answer: P = 13, E = 14, C = 4, and V = 10.* We will also assume you have completed step #3. That is where we will start this example.

STEP 3: *Give all atoms eight electrons. EXCEPT: Hydrogen.*



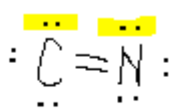
STEP 4: *Check number of electrons, adjust if necessary, and draw brackets to finish.*

A.



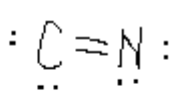
14 electrons are too many. What should we do ? *Add a double bond.*

B.



Again, note the double bond and the "highlighted electrons that will need to be removed.

C.



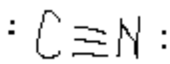
We now have 12 electrons. Still too many. What can we do ? *Add a triple bond.*

D.

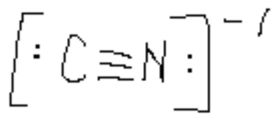
We now have a triple bond, and two more pairs of electrons that must be removed.

E.

This time we are left with 10 electrons. The "4 Key Numbers" calculated earlier determined this was the correct amount. Just one more step.



F.



All finished.

The Mole Concept.

5) **Mole -** the amount of substance that contains Avagadro's number of particles of that substance. Abbreviated: mol. **NOT m, M, or M.**

6) **Avagadro's number -** this is equal to $6.02 * 10^{23}$.
   a) It is specifically the number of C^{12} atoms in 12 grams (the chart mass in grams) of this isotope.
   b) Think of this word like the word "Dozen", which always means twelve. Well this number always means $6.02 * 10^{23}$ particles.

7) **Atomic weight -** the weighted average of the masses of an isotope of an element.
   a) Based on C^{12} , 1 atom = 12 AMUs (atomic mass units).
   b) **AMU -** arbitrary unit equal to 1/12 the mass of a C^{12} atom.
      i) Imagine if we had to determine the "weight" of everyone in the classroom, but, we didn't have a scale. We could select someone and then assign to that individual an arbitrary "weight" unit. We might say he/she weight 10 CRUs (Chem. Room Units). Everyone in the room will now be weighed relative to that student. A person who was twice as heavy would weigh how much ? *Answer: 20 CRUs.* Likewise, some one who weighed 80% of the weight of our selected student, would weigh how much ? *Answer: 8 CRUs.* Everyone's weight would be relative to our original student who set the standard as an arbitrary "10 CRUs".
      ii) The atomic weight of all elements is relative to the "weight" of carbon 12.

8) **Gram Atomic Weight -** the mass in grams of 1 mole of a substance.
   a) .Example 1: Hydrogen.
      i) 1 atom weighs ? *Answer: 1.0079 AMU.*

ii) 1 mole of hydrogen atoms weighs ? *Answer: 1.0079 grams.*

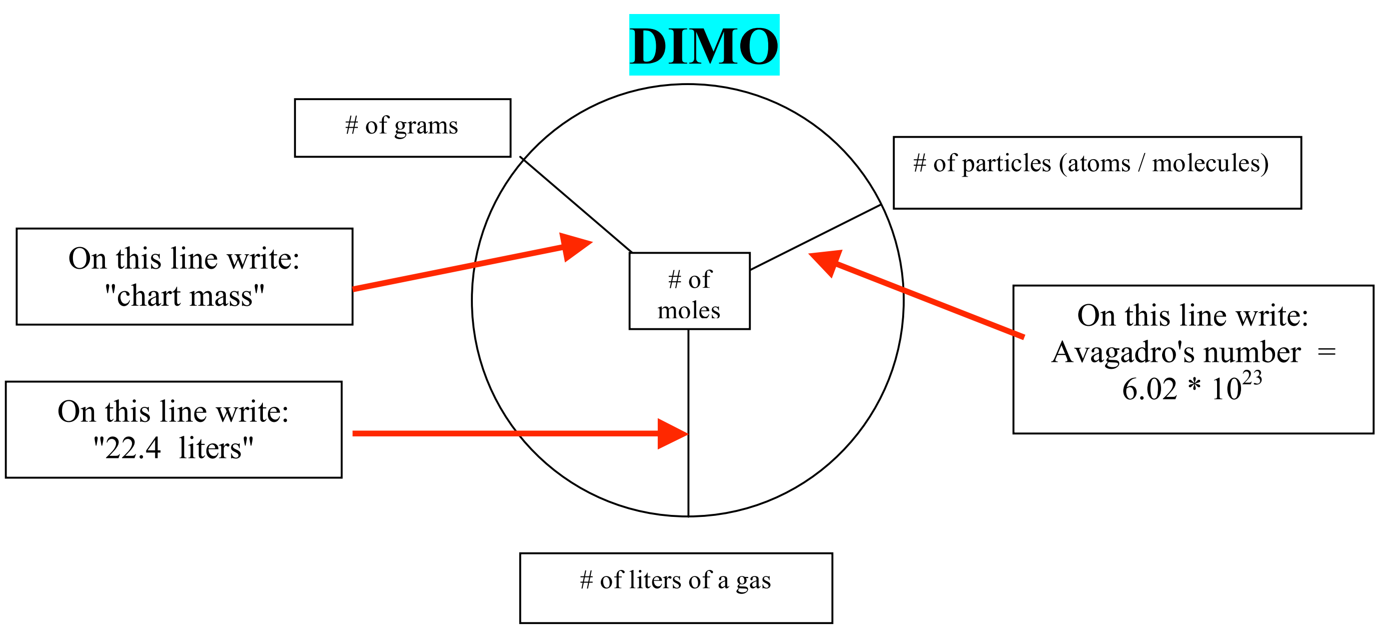
However what is special about Hydrogen ? *Answer: It is a BrINClHOF (diatomic) element.* With that thought in mind:

iii) 1 molecule of Hydrogen weighs ? *Answer: 2.0158 AMUs.*
iv) 1 mole of Hydrogen gas weighs ? *Answer: 2.0158 grams.*

b) Example 2: Water (H_2O)
   i) 1 molecule of water weighs ? *Answer: ~18 AMUs*
   ii) 1 mole of water weighs ? *Answer: ~18 grams.*
c) Important to remember:
   Atoms → Molecules (which may be:) → Diatomic → Polyatomic

9) **Molarity -** is a concentration unit in: #moles / liter. (The liter is usually water)
   a) Abbreviated: M **NOT mol, M, m, or mol.**
   b) Note that the numerator is # of moles. We will learn how to convert this to grams. I call the technique "DIMO".
      i) **DIMO -** an acronym that means: Divide In Multiply Out.

         (1) How to draw the DIMO chart.



(2) How to use the DIMO chart.

(a) There are 4 basic steps that will solve any problem dealing with concentration, volume, or mass.

(i) Determine the chart mass of the substance you're working with.

(ii) Deal with the concentration of that substance.

(iii) Deal with the volume of the solution you're working with.

(iv) Use "DIMO" to solve.

(b) Example 1:

How many grams of NaOH are in 1 liter of a 1<u>M</u> solution of the NaOH?

(i) STEP 1: Determine the chart mass of the substance you're working with. What is the chart mass of NaOH? *Answer: 40 grams. (23 g. of Na, 16 g. of O, 1 g. of H).*

(ii) STEP 2: Deal with the concentration of that substance. 1<u>M</u> = what? *Answer: 1<u>M</u> = 1 mol / liter.*

(iii) STEP 3: Deal with the volume of the solution you're working with. (Set up a proportion using the previous steps info.)

1 mol / liter = x mol / (use the volume given in the problem here) 1 liter.

Therefore "x" = 1 mol.

(iv) STEP 4: Use "DIMO" to solve. (Using the DIMO chart, place the calculated value of moles (1 in

this example) and since you're trying to find "grams", you will multiply out (of the DIMO circle) by multiplying by the "chart mass".) How many grams of NaOH are in 1 liter of a 1M solution of NaOH? *Answer: 40 grams.*

(c) Example 2:
0.75 liters of a 0.5M solution contains how many grams of $CaCl_2$?

(i) STEP 1: Determine the chart mass of the substance you're working with. What is the chart mass of $CaCl_2$? *Answer: 111 grams. (40 g. of Ca, 35.5 g. / atom of Cl, but you have 2 atoms of Cl; therefore you have 71 grams of Cl).*

(ii) STEP 2: Deal with the concentration of that substance. 0.5M = what? *Answer: 0.5M = 0.5 mol / liter.*

(iii) STEP 3: Deal with the volume of the solution you're working with. (Set up a proportion using the previous steps info.)

0.5 mol / liter = x mol / (use the volume given in the problem here) 0.75 liter.
Therefore "x" = 0.375 mol.

(iv) STEP 4: Use "DIMO" to solve. (Using the DIMO chart, place the calculated value of moles (0.375 in this example) and since you're trying to find "grams", you will multiply out (of the DIMO circle) by multiplying by the "chart mass".) How many grams of $CaCl_2$ are in 0.75 liters of a 0.5M solution of $CaCl_2$? *Answer: 41.625 grams.*

Notice in the previous examples we calculated the number of grams of a substance. In the next example, we will determine how many moles of a substance are in a given volume of a solution.

(d) Example 3:

What is the Molarity of a solution created by dissolving 150 grams of NaI into 250 mL of distilled water?

(i) STEP 1: Determine the chart mass of the substance you're working with. What is the chart mass of NaI? *Answer: 150 grams. (23 g. of Na, 127 g. of I).*

(ii) STEP 2: Deal with the concentration of that substance. In this example there is not a concentration given directly. However, we do know that 150 g. NaI / 0.25 Liters. (How did we convert 250mL into 0.25 L? *Answer: We moved the decimal 3 places to the left. (If you don't remember how , when or why to do this: go back to unit #1 and review topic #15; "The Conversion Line".)*

(iii) STEP 3: Deal with the volume of the solution you're working with. (Set up a proportion using the previous steps info.)

150 g. / 0.25 liter = x g. / (since we are looking for Molarity (# mol / Liter) we will use 1 liter here.) 1 Liter

Therefore: 150 g. / 0.25 L = "x" g / 1 Liter.

"x" = 600 grams.

(iv) STEP 4: Use "DIMO" to solve. (Using the DIMO chart, place the calculated value of grams (600 g. in this example) and since you're trying to find "moles", you will divide in (to the DIMO circle) by takings grams calculated divided by the "chart mass".) How many moles of NaI are in 1 liters of the above solution? *Answer: 4 moles.* Therefore, what is the Molarity of this solution? *Answer: 4<u>M</u> (pronounced "Four molar").*

(e) Finding the number of particles (atoms or molecules) is done the same way: except you would divide or multiply in or out by "Avagadro's number". You will get practice doing this on your study guide.

10) **Electrochemistry -** a branch of chemistry dealing with the relations between chemical energy and electrical energy.
   a) Examples of technology utilizing electrochemistry:
      i) Batteries
      ii) Electroplating

11) **Oxidation -** chemical process where electrons are lost. The oxidation number (charge) gets more positive.
12) **Reduction -** chemical process where electrons are gained. The oxidation state (charge) gets more negative.

13) **OIL RIG -** an acronym to help remember the difference between oxidation and reduction. Oxidation Is Loss, Reduction Is Gain (of electrons).

14) **Electrochemical cell -** a system consisting of electrodes that dip into an electrolytic solution and in which a chemical reaction produces an electric current.

   a) **Electrolytic solution -** an aqueous solution that contains ions (electrolytes) and conducts electric current.
      i) Examples:
         (1) "Gatorade", "All-sport" …
         (2) "Battery acid" (although it is a base in alkaline batteries).
         (3) Your blood and interstitial fluid.

   b) There are two types of electrochemical cells:
      i) **Voltaic cell -** uses energy from a spontaneous chemical reaction to produce an electric current. These cells are often called a "Galvanic cells".
         (1) Examples:
            (a) A non-rechargeable battery ("Dry cell")
            (b) A car battery
            (c) A potato clock
            (d) A "plating solution and cell" that does not require additional electricity. (We will copper plate tin this way in the lab.)
         (2) **Anode -** is always the electrode where oxidation occurs. It is the negative terminal in a voltaic cell.
         (3) **Cathode -** is always the electrode where reduction occurs. It is the positive terminal in a voltaic cell.
         (4) Diagrams on following pages.

ii) **<u>Electrolytic cell -</u>** uses energy from a non-spontaneous chemical reaction to produce an electric current. These are often used to plate metals.

(1) Examples:

(a) Electrolysis of water into Hydrogen and Oxygen.

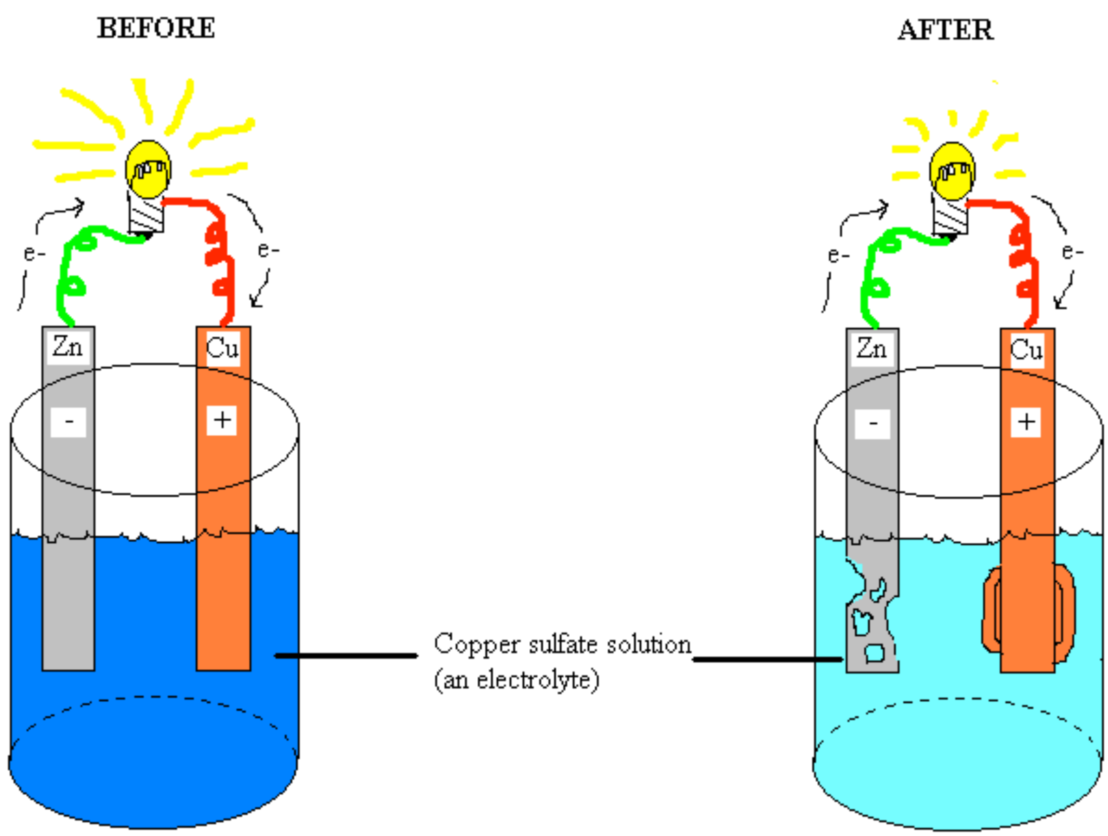
(b) Electrolysis of brine (concentrated salt water) into Chlorine gas, Hydrogen, and Sodium chloride.

(2) **<u>Anode -</u>** <u>is always</u> the electrode where oxidation occurs. It is the positive terminal in an electrolytic cell.

(3) **<u>Cathode -</u>** <u>is always</u> the electrode where reduction occurs. It is the negative terminal in an electrolytic cell.

(4) Diagrams on following pages.

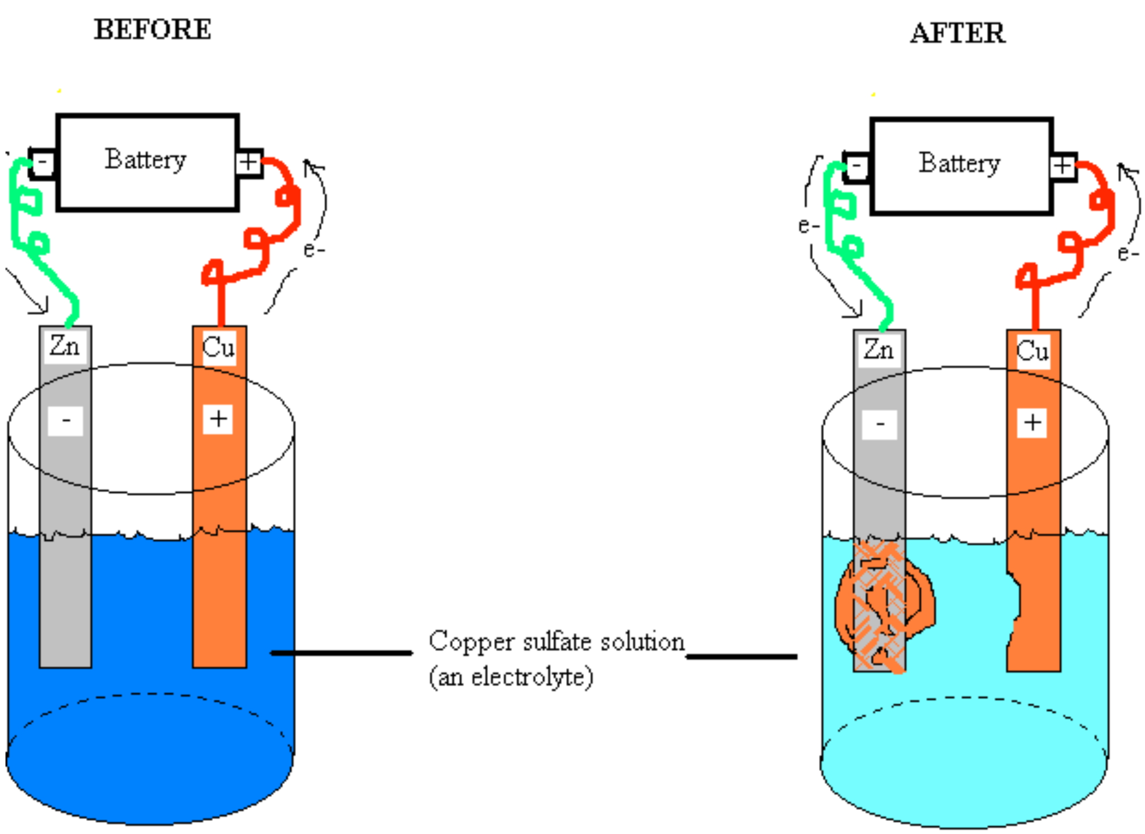
Diagram of a voltaic cell.



Notice that as time goes on the copper will spontaneously plate on to the copper cathode. Simultaneously, the zinc (which is loosing

electrons and putting Zinc cations into solution) will become severely corroded, causing the voltage to decrease and the light to dim.   Copper when in solution creates a blue liquid. As the copper leaves the solution by plating onto the copper strip the color fades.   This is essentially what happens when a battery "dies".

Diagram of an electrolytic cell.



Notice how the external power source forces the electrons to travel opposite their spontaneous direction.   This time the copper is oxidized and thus corroded (albeit after the majority of copper cations in solution have plated the tin).   The tin is plated with the copper.   Once the external power source is removed and the "light" is replaced, this cell will once again become a voltaic cell.

15) **Half-cell reaction -** the chemical reaction that occurs at one of the electrodes in an electrolytic cell. It involves the electrode and electrons only.

a) If oxidation is occurring, the electrode is losing electrons.

i) $Cu \rightarrow\rightarrow\rightarrow Cu^{+2} + 2e^{-}$

(1) In this example copper metal will be converted into copper ions. (Often a blue liquid)

b) If reduction is occurring, the electrode is gaining electrons.

i) Zn+2 + 2e- $\rightarrow\rightarrow\rightarrow$ Zn

(1) In this example zinc ions in solution will be converted into zinc metal.

16) **Half-cell potential -** the sum of the voltages (both positive and / or negative from the oxidation and reduction reactions that occur in an electrolytic cell.

a) **Standard reduction potential -** a measure of the voltage gained or lost when an electrode is reduced in an electrochemical cell.

b) **Cell potential -** the sum of the standard reduction potential and the standard oxidation potential (S.R.P. with the opposite charge.)

i) Cell potentials are always calculated from the standard reduction potential.

ii) The electrode that is being oxidized will have the same value, but the opposite charge; of what is listed as the S.R.P..

iii) Example: 1 Using the metals pictured in the previous diagrams: Cu & Zn.

Calculate the cell potential based on their standard reduction potentials.

(1) STEP 1: Write down the S.R.P. of the metals involved. (You can find these listed in a table in any chemistry text book.) For these elements they are:

$Cu^{+2} + 2e^{-} \rightarrow\rightarrow\rightarrow Cu \qquad = +0.34$ volt
$Zn^{+2} + 2e^{-} \rightarrow\rightarrow\rightarrow Zn \qquad = -0.76$ volt

(2) STEP 2:

(a) Rewrite the more positive voltage and the corresponding chemical equation. (Since it is the most positive, we will not change it; and it will be the electrode where *Reduction* occurs.)

$Cu^{+2} + 2e^{-} \rightarrow\rightarrow\rightarrow Cu \qquad = +0.34$ volt

(b) Write the opposite of the remaining voltage and its corresponding chemical equation. (For the voltage, all you have to do is switch the "sign" of the charge. For the equation: you will switch everything from the left side of the arrow to the right side, and everything from the right side to the left side.)

$Zn \rightarrow\rightarrow\rightarrow Zn^{+2} + 2e^{-} \qquad = +0.76$ volt

(3) STEP 3:

Determine if the electrons on both sides of the equation cancel each other out. If they do not, you will have to multiply one or both of the half-cell reactions by a whole number until they do cancel. (What ever number you choose to multiply the reaction equation by must also be multiplied with the voltage.) As you will see they do cancel in this example.

$Cu^{+2} + \mathbf{2e^{-}} \rightarrow\rightarrow\rightarrow Cu \qquad = +0.34$ volt
$Zn \rightarrow\rightarrow\rightarrow Zn^{+2} + \mathbf{2e^{-}} \qquad = +0.76$ volt

(4) STEP 3:

Calculate the cell potential by adding the voltages determined above together. (This is how much voltage you should get from the two metals in a 1<u>M</u> solution of electrolyte)

(+0.34 volt)
+(+0.76 volt)
+1.10 volt = cell potential **(Since this is a positive value this would be a spontaneous reaction, and thus a voltaic cell.)**

iv) Example 2: Using the metals Zn^{+2}/Zn & Cr^{+3}/Cr Calculate the cell potential based on their standard reduction potentials.

(1) STEP 1: Write down the S.R.P. of the metals involved. (You can find these listed in a table in any chemistry text book.) For these elements they are:

$Cr^{+3} + 3e^{-} \longrightarrow\longrightarrow\longrightarrow Cr \quad = -0.74$ volt
$Zn^{+2} + 2e^{-} \longrightarrow\longrightarrow\longrightarrow Zn \quad = -0.76$ volt

(2) STEP 2:

(a) Rewrite the more positive voltage and the corresponding chemical equation. (Since it is the most positive, we will not change it; and it will be the electrode where *Reduction* occurs.)

$Cr^{+3} + 3e^{-} \longrightarrow\longrightarrow\longrightarrow Cr \quad = -0.74$ volt

(b) Write the opposite of the remaining voltage and its corresponding chemical equation. (For the voltage, all you have to do is switch the "sign" of the charge. For the equation: you will switch everything from the left side of the arrow to the right side, and everything from the right side to the left side.)

$Zn \longrightarrow\longrightarrow\longrightarrow Zn^{+2} + 2e^{-} \quad = +0.76$ volt

(3) STEP 3:
Determine if the electrons on both sides of the equation cancel each other out. If they do not, you will have to multiply one or both of the half-cell reactions by a whole number until they do cancel. (What ever number you choose to multiply the reaction equation by must also be multiplied with the voltage.) As you should notice, they do not cancel in this example.

Cr^{+3} + **$3e^-$** →→→ Cr = -0.74 volt
Zn →→→ Zn^{+2} + **$2e^-$** = +0.76 volt

The smallest common denominator is 6 in this example. We will have to multiply the Chromium by 2; and the Zinc by 3.

2Cr^{+3} + **$6e^-$** →→→ **2**Cr = **-1.48** volt
3Zn →→→ **3**Zn^{+2} + **$6e^-$** = **+2.28** volt

(4) STEP 3:
Calculate the cell potential by adding the voltages determined above together. (This is how much voltage you should get from the two metals in a 1<u>M</u> solution of electrolyte)

(-1.48 volt)
<u>+(+2.28 volt)</u>
+0.80 volt = cell potential **(Again, since this is a positive value this would be a spontaneous reaction, and thus a voltaic cell. If the calculated value would be negative, this cell would not be spontaneous and the reaction would require an external power source. Thus, if negative, it would be electrolytic.)**